$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$; $K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$

Now that we have Ka, solve as a normal weak acid.

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+x	x
NH_3	0	+x	x
NH_4^+	0.100	−x	0.100−x

Let "x" equal the change in hydronium ion concentration

$$\frac{(x)(x)}{0.100 - x} = 5.56 \times 10^{-10}$$

x << 0.100, so

$0.100 - x \simeq 0.100$

$$\frac{x^2}{0.100} = 5.56 \times 10^{-10}$$

$x = 7.45 \times 10^{-6} = [H_3O^+]$

$pH = -\log_{10}(7.45 \times 10^{-6}) = \boxed{5.12}$

Compare:
pH of 0.100 M nitric acid: 1.00
pH of 0.100 M nitrous acid: 2.17
pH of DI water: 7.00

0.100 M $NaC_2H_3O_2$, Find pH

$$NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O_2^-$$

Na^+: Not a proton donor (acid), since no H+ to donate. Also not likely to be basic, as the positive charge should repel H+. NEUTRAL.

$C_2H_3O_2^-$: Negative charge ... this might accept protons and be a base. If so, the reaction looks like this:

$$C_2H_3O_2^- + H_2O \rightleftharpoons \underline{HC_2H_3O_2} + OH^-$$

Stable in water? This is ACETIC ACID, a common WEAK ACID. Weak acids are stable in water, since very few weak acid molecules ionize!
So, acetate ion is a base.

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^- \; ; \; K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = ?$$

Appendix I doesn't list a Kb for acetate ion, but Appendix H has a Ka for acetic acid, its conjugate.

$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$K_{a,\,HC_2H_3O_2} = 1.8 \times 10^{-5}$$

$$(1.8 \times 10^{-5})\, K_b = 1.0 \times 10^{-14}$$

$$K_b = 5.56 \times 10^{-10}$$

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^- \; ; \; K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = 5.56\times10^{-10}$$

Once we find Kb, just solve like any other weak base.

Species	[Initial]	Δ	[Equilibrium]
OH^-	0	$+x$	x
$HC_2H_3O_2$	0	$+x$	x
$C_2H_3O_2^-$	0.100	$-x$	$0.100-x$

Let "x" equal the change in hydroxide ion concentration!

$$\frac{(x)(x)}{(0.100-x)} = 5.56\times10^{-10}$$

$0.100 - x \approx 0.100$ $(x \ll 0.100)$

$$\frac{x^2}{0.100} = 5.56\times10^{-10}$$

$$x = 7.45\times10^{-5}\,M = [OH^-]$$

$pOH = -\log_{10}(7.45\times10^{-5}) = 4.12$

$pH + pOH = 14.00$

$pH = 14.00 - 4.12 = \boxed{9.87}$

For comparison:

0.100 M sodium acetate, pH = 9.87

0.100 M ammonia, pH = 11.13

0.100 M NaOH (strong base), pH = 13.00

The acetate ion is basic, but it's a very weak base!

0.100 M NaCl, Find pH

$$NaCl \rightarrow Na^+ + Cl^-$$

Na^+: Not a proton donor (acid), since no H+ to donate. Also not likely to be basic, as the positive charge should repel H+. NEUTRAL.

Cl^-; Not a proton donor (acid), since it has no protons (H+) to donate. It does have a negative charge, and may attract protons. If so, this would be the reaction:

$$Cl^- + H_2O \rightleftharpoons \underline{HCl} + OH^-$$

This is hydrochloric acid, a STRONG ACID. Strong acids completely ionize, so they are NOT stable in water. This reaction won't happen! Chloride ion should be NEUTRAL.

In a NaCl solution, water itself controls the pH. So, the pH will be 7.00

POLYPROTIC ACIDS

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?

① $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

② $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$

③ $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$

Phosphoric acid has THREE acidic protons!

$K_{a1} = 7.5 \times 10^{-3}$

$K_{a2} = 6.2 \times 10^{-8}$

$K_{a3} = 4.2 \times 10^{-13}$

The first dissocation is dominant here, and for simple calculations of phosphoric acid in water, we will simply use the first ionization and ignore the other two.

Remember: This is a weak acid. It exists in water mostly as undissociated phosphoric acid molecules.

Solve the equilibrium of phosphoric acid's FIRST proton:

$$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+ ; \; K_a = 7.5 \times 10^{-3}$$

$$K_a = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]} = 7.5 \times 10^{-3}$$

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	$+x$	x
$H_2PO_4^-$	0	$+x$	x
H_3PO_4	0.10	$-x$	$0.10 - x$

Let "x' equal the change in hydronium ion concentration

$$\frac{(x)(x)}{(0.10 - x)} = 7.5 \times 10^{-3}$$

$x << 0.10$

so $0.10 - x \approx 0.10$

$$\frac{x^2}{0.10} = 7.5 \times 10^{-3}$$

$x = 0.0273861279 = [H_3O^+]$

$$\boxed{pH = 1.56}$$

Check this in experiment 16A. You'll measure the pH of this same concentration of phosphoric acid yourself. It may be slightly lower, since we ignored the other two acid ionizations.

143 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

$NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$ Is this salt acidic, basic, or neutral?

NH_4^+: Has protons? Might be a proton donor (acid):

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

Ammonia. A WEAK BASE! This is stable in water, so this reaction will go. Ammonium ion is ACIDIC.

NO_3^-: Could this be a proton acceptor (base)?

$NO_3^- + H_2O \rightleftharpoons HNO_3 + OH^-$

Nitric acid, a STRONG ACID. Not stable in water, so this reaction doesn't go! Nitrate is NEUTRAL.

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$; $K_a = \dfrac{[NH_3][H_3O^+]}{[NH_4^+]} = ?$

Appendix I has Kb for ammonia, so we can calculate Ka for ammonium ion:

$K_{b,NH_3} = 1.8 \times 10^{-5}$ $\quad K_a \times K_b = 1.0 \times 10^{-14}$

$K_a(1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$; $K_a = 5.56 \times 10^{-10}$

We'll need to know the MOLAR concentration of the solution. Calculate.

$3.00\,g\,NH_4NO_3 \times \dfrac{mol\,NH_4NO_3}{80.052\,g\,NH_4NO_3} = 0.03747564 08\,mol\,NH_4NO_3$

$M = \dfrac{0.0374756408\,mol\,NH_4NO_3}{0.250\,L} = 0.1499025633\,M\,NH_4NO_3$

$\uparrow$ 250. mL

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$; $K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$

0.1499025633 M NH_4NO_3

Solve the acid ionization equilibrium.

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+x	x
NH_3	0	+x	x
NH_4^+	0.1499025633	−x	0.1499025633 − x

Let "x" equal the change in hydronium ion concentration

$$\frac{(x)(x)}{(0.1499025633 - x)} = 5.56 \times 10^{-10}$$

$x << 0.150$, so
$0.150 - x \approx 0.150$

$$\frac{x^2}{0.1499025633} = 5.56 \times 10^{-10}$$

$$x = 9.12939 \times 10^{-6} = [H_3O^+]$$

$$\boxed{pH = 5.04}$$